William Thomson, known as Lord Kelvin (1824–1907), first suggested a temperature scale that did not use negative numbers. Kelvin's scale, now adopted as the SI standard, uses the same size degree as the Celsius scale, but it begins at absolute zero. Thus, the freezing point of water is reached at 273.15 degrees above the starting point; that is, 0°C is the same as 273.15 kelvins or 273.15 K. Temperatures on the Celsius scale are readily converted to the kelvin scale, and vice versa, using the relation

\[ T (K) = t (^\circ C) + 273.15 \]

**EXERCISE 1.6  TEMPERATURE CONVERSIONS**

Carry out the following temperature conversions: (a) 25°C to K. (b) Liquified nitrogen boils at 77 K. What is this temperature in °C? In °F?

---

### 1.5 HANDLING NUMBERS

#### PRECISION AND ACCURACY

The **precision** of a measurement refers to the agreement of repeated measurements of a value with one another. **Accuracy** is the agreement between the measured quantity and the accepted value. A highly precise number may be inaccurate because of an error that is the same for each measurement. As an example, consider the data in Table 1.4. Three students were asked to determine the mass of a piece of metal whose mass is known to be 0.520 g.

The data for Student A are neither very precise nor accurate; the individual values differ widely from one another, and the average value is not accurate. Student B was able to determine the mass of the metal more precisely; the three values deviate but little from one another, but the average mass is still not accurate. In contrast, the data for student C is both precise and accurate.

#### SIGNIFICANT FIGURES

Suppose you wish to find the percentage of mass lost (as water) by a popcorn kernel when popped. The mass before heating was found to be 0.123 g, and the mass after heating was 0.108 g. The typical analytical balance, however, is accurate to within 1 mg or 0.001 g. This means that the mass may have been as low as 0.122 g or as high as 0.124 g. The

**TABLE 1.4  Data to Illustrate Precision and Accuracy**

<table>
<thead>
<tr>
<th>MEASUREMENT (g)</th>
<th>1</th>
<th>2</th>
<th>3</th>
<th>AVERAGE (g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Student A</td>
<td>0.521</td>
<td>0.515</td>
<td>0.509</td>
<td>0.515</td>
</tr>
<tr>
<td>Student B</td>
<td>0.516</td>
<td>0.515</td>
<td>0.514</td>
<td>0.515</td>
</tr>
<tr>
<td>Student C</td>
<td>0.521</td>
<td>0.520</td>
<td>0.520</td>
<td>0.520</td>
</tr>
</tbody>
</table>

**FIGURE 1.12**

implied error is 1 part in 123 or about 0.8%. Similarly, writing 0.108 g implies the mass is between 0.107 or 0.109 g. Both numbers you observed had three significant figures; all three digits are important and were experimentally determined.

The % mass loss can be calculated as

\[
\text{\% Mass lost} = \frac{\text{mass of water lost}}{\text{mass of kernel}} \times 100 = \frac{(0.123 - 0.108)g}{0.123 \text{ g}} \times 100
\]

When performing this calculation on an electronic calculator, you would observe 12.19512195. An answer with this many digits implies that you made the original measurement far more precisely than was actually the case. The 10-digit number implies that the final digit is 4, 5, or 6 and that all preceding digits are correct. That is, it implies you are sure of the result to 1 part in 1.2 billion. This is certainly not the case! Instead, the proper result is 12%, a number with two significant figures.

The following guidelines regarding significant figures are helpful, but above all else, use common sense!

**GUIDELINES FOR DETERMINING SIGNIFICANT FIGURES**

**Rule 1.** To determine the number of significant figures in a measurement, read the number from left to right and count all of the digits, starting with the first digit that is not zero.

<table>
<thead>
<tr>
<th>EXAMPLE</th>
<th>NUMBER OF SIGNIFICANT FIGURES</th>
</tr>
</thead>
<tbody>
<tr>
<td>1.23 g</td>
<td>3</td>
</tr>
<tr>
<td>0.00123 g</td>
<td>3</td>
</tr>
<tr>
<td>2.0 g and 0.020 g</td>
<td>Both of these numbers have two significant digits. When a number is greater than 1, all zeros to the right of the decimal point are significant. For a number less than 1, all zeros to the right of the first significant digit are significant.</td>
</tr>
<tr>
<td>100 g</td>
<td>In numbers that do not contain a decimal point, &quot;trailing&quot; zeros may or may not be significant. If only the last digit is uncertain there are three significant figures, so the number should be written in scientific notation as (1.00 \times 10^2). Alternatively, this idea can be conveyed by following the number with a decimal point (100.), a practice followed in this book. It is also possible that there is an error of 10 in the number, in which case there are only two significant figures, and the number should be written as (1.0 \times 10^2). Finally (1 \times 10^2) would imply only one significant figure.</td>
</tr>
<tr>
<td>100 cm/meter</td>
<td>Infinite number of significant figures, because this is a defined quantity.</td>
</tr>
<tr>
<td>(\pi = 3.1415926\ldots)</td>
<td>The value for (\pi) is a known to a large number of significant figures; you may choose the number of digits appropriate to your calculation.</td>
</tr>
</tbody>
</table>

**Rule 2.** When adding or subtracting, the number of decimal places in the answer should be equal to the number of decimal places in the number with the fewest places.
1.6 PROBLEM SOLVING BY DIMENSIONAL ANALYSIS

0.12 2 significant figures 2 decimal places
1.6 2 significant figures 1 decimal place
10.967 5 significant figures 3 decimal places
12.696

The answer above should be reported as 12.7, a number with 1 decimal place, because 1.6 has only 1 decimal place.

Rule 3. In multiplication and division, the number of significant figures in the answer should be the same as that in the quantity with the fewest significant figures.

\[
\frac{0.01208}{0.0236} = 0.512 \text{ or, in exponential notation, } 5.12 \times 10^{-1}
\]

Since 0.0236 has only three significant figures, the answer must be limited to three significant figures.

Rule 4. When a number is rounded off (that is, the number of significant figures is reduced), the last digit retained is increased by 1 only if the first digit to be dropped is 5 or greater.\(^*\)

<table>
<thead>
<tr>
<th>FULL NUMBER</th>
<th>NUMBER ROUNDED TO THREE SIGNIFICANT FIGURES</th>
</tr>
</thead>
<tbody>
<tr>
<td>12.696</td>
<td>12.7</td>
</tr>
<tr>
<td>16.249</td>
<td>16.2</td>
</tr>
<tr>
<td>18.35</td>
<td>18.4</td>
</tr>
<tr>
<td>18.351</td>
<td>18.4</td>
</tr>
</tbody>
</table>

One last word regarding significant figures and calculations. In working problems in chemistry on a pocket calculator, you should carry through the calculation all of the digits allowed by the calculator and round off only at the end of the problem. Rounding off in the middle can introduce significant errors. If your answers do not quite agree with those in the back of the book, this may be the source of the disagreement.

EXERCISE 1.7 SIGNIFICANT FIGURES
(a) How many significant figures are there in 12.63 and in 0.063? (b) What is the sum of 12.63 and 0.063? (c) What is the product of 12.63 and 0.063?

1.6 PROBLEM SOLVING BY DIMENSIONAL ANALYSIS

Dimensional analysis is a systematic way of solving numerical problems. Simply put, every number in a problem must have some units associated with it; density, for example, is given as mass per unit volume (g/cm\(^3\)). When the numbers in a calculation are manipulated in the correct way, their units must cancel out to leave the final answer in the appropriate units. If you set up the problem incorrectly, the units will not cancel properly, and you will know immediately that you have made a mistake.

Dimensional analysis has already been used in the examples in this chapter. However, since the method is such a valuable tool in chemistry,

\(^*\)A modification of this rule is sometimes used to reduce the accumulation of roundoff errors: If the first digit dropped is 5 and there are no following digits or all following digits are zeros (e.g., 18.35 or 18.3500), then the last digit retained is increased by 1 only if it is odd. Thus, 18.35 and 18.45 are both rounded to 18.4.
it is worthwhile to illustrate it further using some more familiar items and the type of calculation you have often done in your head. Let us say you need 30 cans of soft drinks for a party, and the cost is $2.15 per six-pack of cans. What is the total cost? The strategy in solving the problem is first to convert the units "cans" to the unit "six-packs," since the cost is given in units of "dollars per six-pack."

**Step 1.** Find the number of six-packs required. Convert units of "cans" to units of "six-packs."

\[
\text{Number of six-packs} = 30 \text{ cans} \times \frac{1 \text{ six-pack}}{6 \text{ cans}} = 5 \text{ six-packs}
\]

Notice that if you had multiplied 30 times 6, the units would have come out to be (cans²/six-pack), which is clearly nonsense.

**Step 2.** Find the total cost. Convert units of "six-packs" into units of "$."

\[
\text{Total price} = 5 \text{ six-packs} \times \frac{2.15 \text{ $}}{\text{six-pack}} = 10.75 \text{ $}
\]

Now that you have worked through the example, notice that in each step you always multiplied the starting unit by a factor that gave the answer in the desired unit. Such factors as (1 six-pack/6 cans) or ($2.15/six-pack) are called conversion factors. Conversion factors are multipliers that relate the desired unit to the starting unit.

Conversion factor = \( \frac{\text{desired unit}}{\text{starting unit}} \)

The numerator of the factor must be equal or equivalent to the denominator. Thus, 1 six-pack is equivalent to 6 cans, and $2.15 is the amount you must pay for 1 six-pack. Other "conversion factors" are 453.6 g/pound, 100 cm/m, or 5°C/9°F.

**EXAMPLE 1.7**

**DIMENSIONAL ANALYSIS: DENSITY AND VOLUME**

If a frying pan needs a Teflon coating that is 1.00 mm thick, and the area covered is 36.0 square inches, how many pounds of Teflon are required to coat the pan? Teflon has a density of 0.805 g/cm³.

**Solution** The first step of the problem is to calculate the total volume of Teflon needed. Once this is done, you can convert the volume to a mass by using the density.

**Step 1. Volume of Teflon required.**

The volume of the Teflon can be calculated from the product of the thickness and the area. Unfortunately, these dimensions are in different units and cannot be directly multiplied until both are in the same unit. You could convert the thickness to inches, so that the product of thickness (inches) and area (inches²) gives the volume in cubic inches (inches³). Alternatively, you could convert the area to square millimeters so that volume is obtained in cubic millimeters. However, we choose to do neither. Thinking ahead to the next step in the problem, you will want to use the
density in units of g/cm³. Therefore, it would be most useful to obtain the volume of Teflon in units of cm³ so that it is compatible with density. Thus,

\[
\text{Thickness} = 1.00 \text{ mm} \left( \frac{1 \text{ cm}}{10 \text{ mm}} \right) = 0.100 \text{ cm}
\]

\[
\text{Area} = 36.0 \text{ in}^2 \left( \frac{2.54 \text{ cm}}{	ext{in}} \right)^2 = 232 \text{ cm}^2
\]

The factor 2.54 cm/in is given in the table at the back of the book. Since you need cm²/in², you simply square the conversion factor and its units (= 6.45 cm²/in²).

With the area and thickness in the same units, you can now calculate the volume.

\[
\text{Volume of Teflon coating} = \text{thickness} \times \text{area} = (0.100 \text{ cm})(232 \text{ cm}^2) = 23.2 \text{ cm}^3
\]

Step 2. Weight of Teflon required.

In solving any problem, you should keep firmly in mind the answer for which you are aiming and, just as importantly, the units of that answer. In this case, you want the weight of Teflon in pounds. As you saw in Example 1.4, the density will allow you first to convert the volume of Teflon to a mass in grams.

\[
23.2 \text{ cm}^3 \left( \frac{0.805 \text{ g}}{\text{cm}^3} \right) = 18.7 \text{ g}
\]

Now that you have the mass in grams, it is a simple matter to convert it to the weight in pounds, since you know (from the table at the back of the book) that 1 pound is equivalent to 454 g.

\[
18.7 \text{ g} \left( \frac{1 \text{ pound}}{454 \text{ g}} \right) = 0.0412 \text{ pounds} = 4.12 \times 10^{-2} \text{ pounds}
\]

Notice that the final answer has three significant figures.

**EXERCISE 1.8 MASS, VOLUME, AND DENSITY**

Mercury is a metal, but unlike almost all other metals, it is a liquid with a density of 13.6 g/cm³ at room temperature. It is also poisonous and should be treated with great care and respect.

(a) If you have 100. mL of mercury, how many grams do you have?

How many pounds?

(b) If you spill 100. mL of mercury on the floor, and it spreads out into a puddle that is 2.0 mm thick, what area of the floor is covered? Calculate the area in square centimeters and square inches. (Incidentally, mercury is also expensive; 100 mL would cost approximately $90)

**SUMMARY**

In science, an investigation of a problem usually follows a well worn pathway (Section 1.2). After reviewing work done before, the scientist constructs a **hypothesis**, a tentative explanation or prediction of the observed facts. This hypothesis is revised as further experimental work is done, and the final result can be a **law**, a concise verbal or mathematical statement of a relation that is always the same under the same conditions. To explain the law, we attempt to devise a **theory**, a principle that explains a body of facts and the laws based on them.